

Chapter 8 - Inorganic carbon chemistry

Limestone

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Limestone

Environmental Problems in the Peak District

Local residents have been protesting against proposals to site a new limestone quarry in the beautiful Derbyshire Peak District. However, limestone is such a useful and sought-after mineral that demand has encouraged mining in National Park areas.

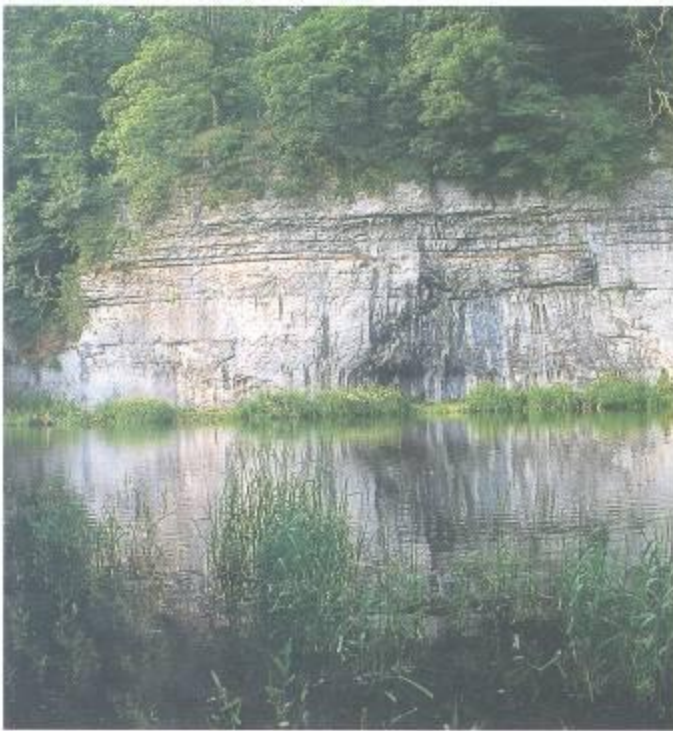


Figure 8.1 This gorge in the Peak District is made from limestone.

As the newspaper article on the left says, limestone is found in the Peak District, an example being this gorge on the River Wye (Figure 8.1). In England, it is also found in the Yorkshire Dales, the Sussex Downs and the Mendip Hills. As well as giving rise to beautiful and varied countryside, limestone is a very useful raw material.

Limestone is composed of calcium carbonate (CaCO_3) in the form of the mineral calcite (Figure 8.2). Chalk and marble are also made of calcite which is the second most abundant

mineral in the Earth's crust after the different types of silicates (which include clay, granite and sandstone).

Chalk is made of the 'shells' of marine algae (that is, plants). It is a form of limestone. Most other limestones are formed from the debris of animal structures, for example brachiopods and crinoids.

Marble is a metamorphic rock made of calcium carbonate. It is formed when limestone is subjected to high pressures or high temperatures, or

sometimes both acting together, to create crystals of calcium carbonate in the rock.

In a typical year, about 80 million tonnes of lime-stone are quarried in Great Britain. Although it is cheap to quarry, as it is found near the surface, there are some environmental costs in its extraction. List the environmental issues which could arise through the quarrying of limestone.

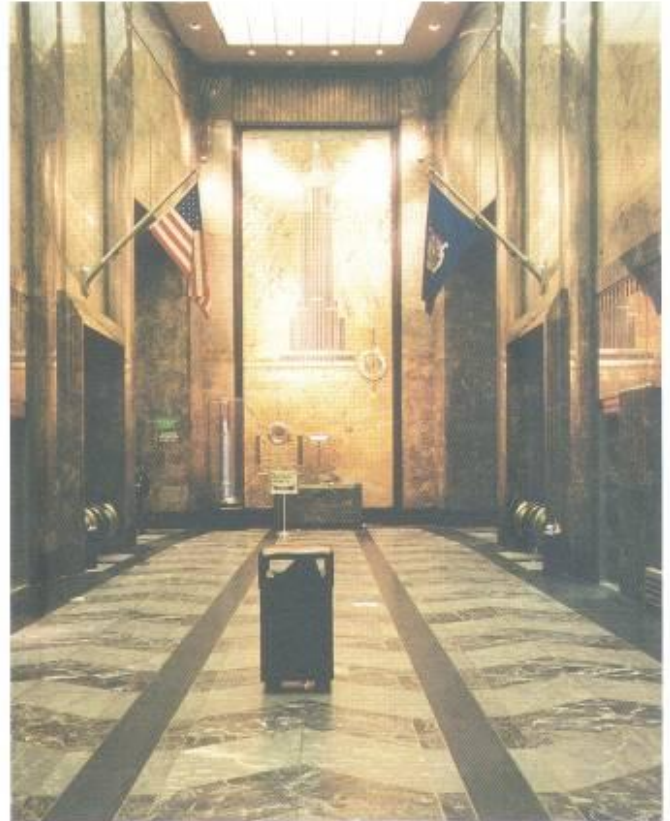


Figure 8.2 Chalk, calcite and marble are all forms of calcium carbonate.

Direct uses of limestone

Limestone has a variety of uses in, for example, the making of cement, road building, glass making and the extraction of iron (Figure 8.3).

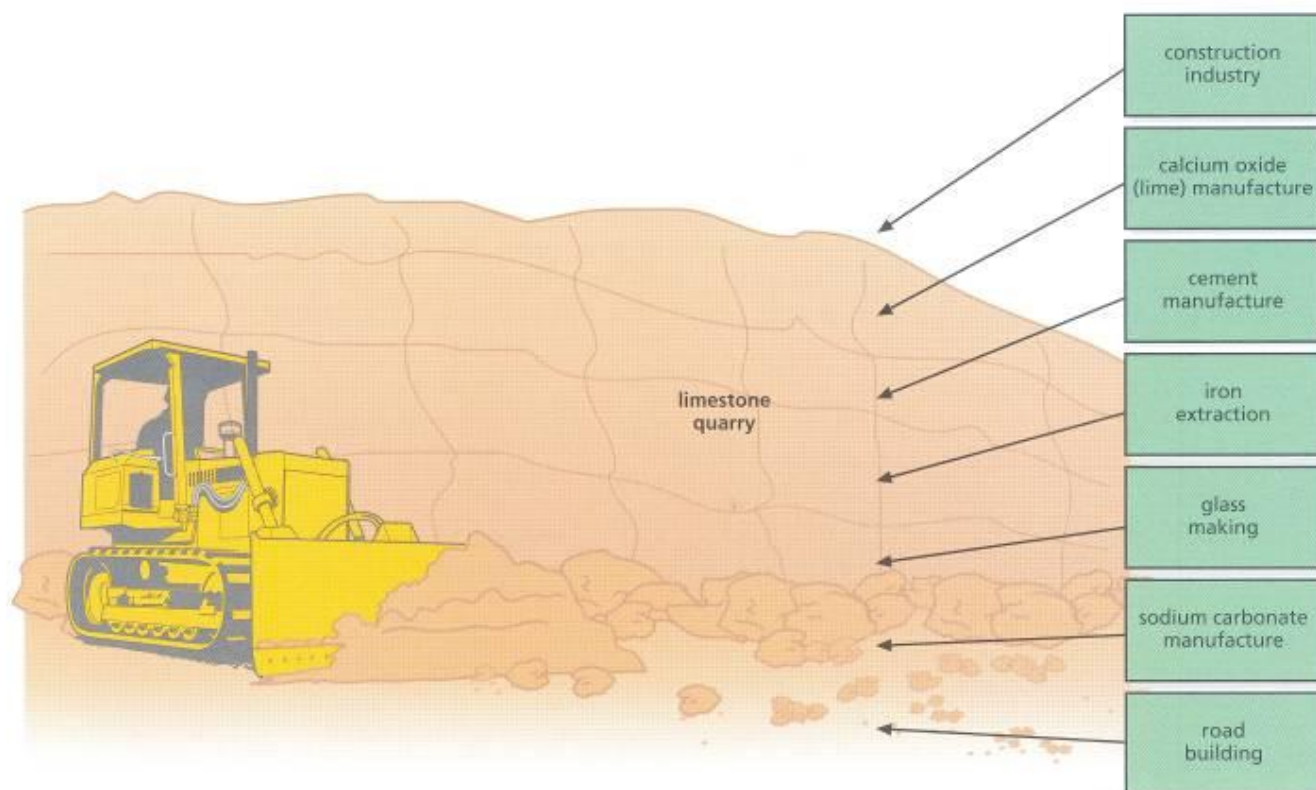


Figure 8.3 Uses of limestone.

Neutralisation of acid soil

Powdered limestone is most often used to neutralise acid soil (Figure 8.4) because it is cheaper than any form of lime (calcium oxide), which has to be produced by heating limestone. The reaction of limestone with acidic soil can be shown by the following ionic equation.

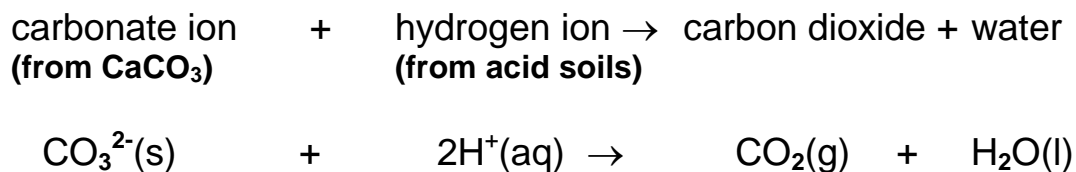




Figure 8.4 Spreading limestone on to soil.

Manufacture of iron and steel

In the blast furnace, limestone is used to remove earthy and sandy materials found in the iron ore. A liquid slag is formed, which is mainly calcium silicate. More details of the extraction of iron and its conversion into steel are given in Chapter 9.

Manufacture of cement and concrete

Limestone (or chalk) is mixed with clay (or shale) in a heated rotary kiln, using coal or oil as the fuel (Figure 8.5). The material produced is called cement. It contains a mixture of calcium aluminate ($\text{Ca}(\text{AlO}_2)_2$) and calcium silicate (CaSiO_3).

The dry product is ground to a powder and then a little calcium sulphate (CaSO_4) is added to slow down the setting rate of the cement. When water is added to the mixture slow complex chemical changes occur, resulting in the formation of a hard interlocking mass of crystals of hydrated calcium aluminate and silicate.



Figure 8.5 A modern rotary kiln.

Concrete is a mixture of cement with stone chippings and sand, which help to give it body. After the ingredients have been mixed with water they are poured into wooden moulds and allowed to set hard. Reinforced concrete is made by allowing concrete to set around steel rods or mesh to give it greater tensile strength, which is required for the construction of large bridges (Figure 8.6) and tall buildings.

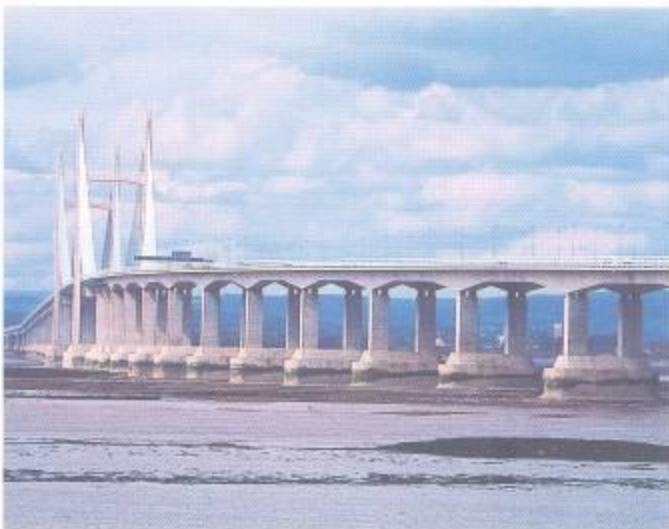


Figure 8.6 The Severn Bridge.

Manufacture of sodium carbonate — the Solvay process

Sodium carbonate (Na_2CO_3) is one of the world's most important industrial chemicals. It is used in the manufacture of soaps, detergents, dyes, drugs and other chemicals. The world production of sodium carbonate now exceeds 30 million tonnes per year. Sodium carbonate is manufactured ingeniously

using calcium carbonate, sodium chloride, carbon dioxide and ammonia in a continuous process, with the Solvay tower the centre of reactions. Sodium carbonate is made as economically as possible because both the carbon dioxide and the ammonia components are recycled continuously.

Indirect uses of limestone

Lime manufacture

When calcium carbonate is heated strongly it thermally dissociates (breaks up reversibly) to form calcium oxide (lime) and carbon dioxide.

calcium carbonate \leftrightarrow calcium oxide + carbon dioxide



This reaction can go in either direction, depending on the temperature and pressure used. This reaction is an important industrial process and takes place in a lime kiln (Figure 8.7). Two million tonnes of calcium oxide are produced in the UK every year.

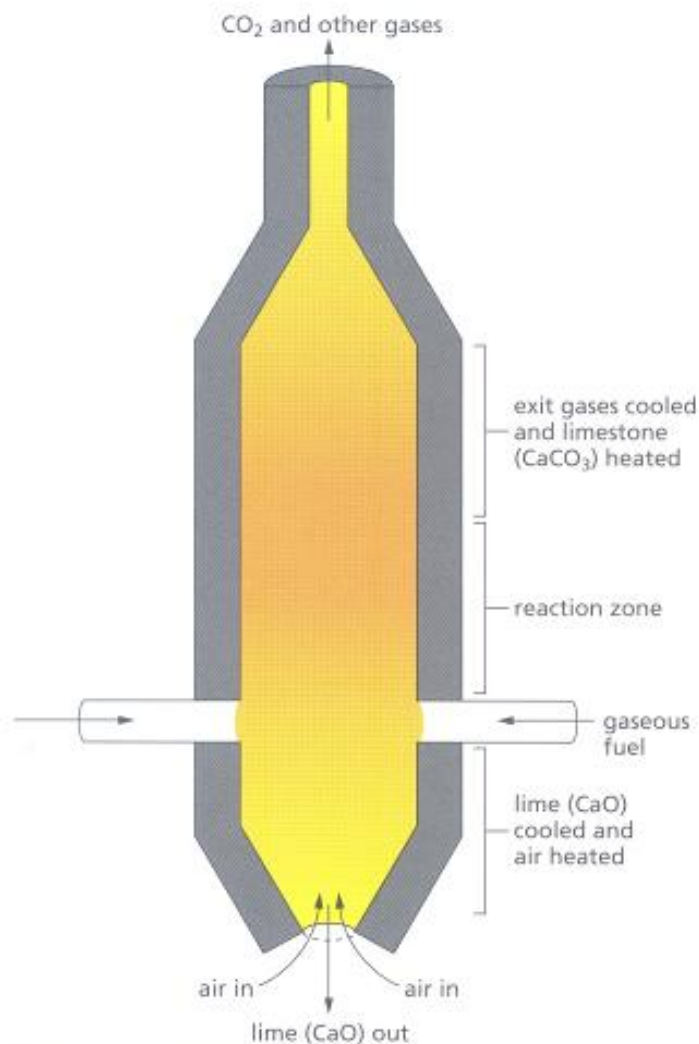


Figure 8.7 The calcium oxide produced from this process is known as quicklime or lime and is used in large quantities in the manufacture of soda glass.

Calcium oxide (CaO) is a base and is still used by some farmers to spread on fields to neutralise soil acidity and to improve drainage of water through soils that contain large amounts of clay. It also has a use as a drying agent in industry. Soda glass is made by heating sand with soda (sodium carbonate, Na_2CO_3) and lime (Figure 8.8). For further discussion of glasses see Chapter 4.

Large amounts of calcium oxide are also converted into calcium hydroxide ($\text{Ca}(\text{OH})_2$) which is called **slaked lime**.

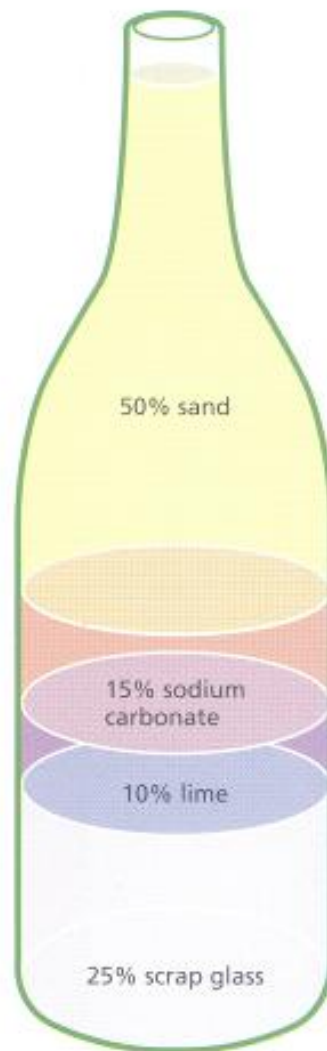


Figure 8.8 The composition of glass.

Manufacture of calcium hydroxide — slaked lime

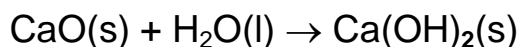
Calcium hydroxide is a cheap industrial alkali (Figure 8.9). It is used in large quantities to make bleaching powder, by some farmers to reduce soil acidity, in the manufacture of whitewash, in glass manufacture and in water purification. Calcium hydroxide, in its white powder form, is produced by adding an equal amount of water to calcium oxide in a carefully controlled reaction. The control is needed because it is a very exothermic reaction.



Figure 8.9 Large-scale production of calcium hydroxide.

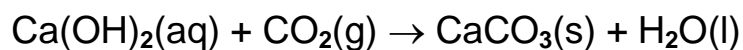
This process can be shown on the small scale in the laboratory by heating a lump of limestone very strongly to convert it to calcium oxide. Water can then be carefully added dropwise to the calcium oxide. An exothermic reaction takes place as the water and calcium oxide react together in this slaking process to form calcium hydroxide.

calcium oxide + water \rightarrow calcium hydroxide



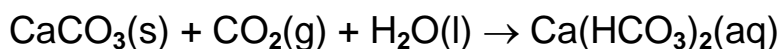
A weak solution of calcium hydroxide in water is called **limewater**. It is used to test for carbon dioxide gas, as a white solid of calcium carbonate is formed if carbon dioxide gas is mixed with it (Figure 8.10):

calcium hydroxide + carbon dioxide \rightarrow calcium carbonate + water



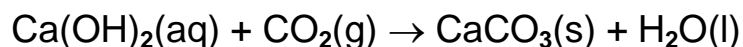
If carbon dioxide is bubbled for a further length of time then the white precipitate of calcium carbonate dissolves and a solution of calcium hydrogencarbonate is produced:

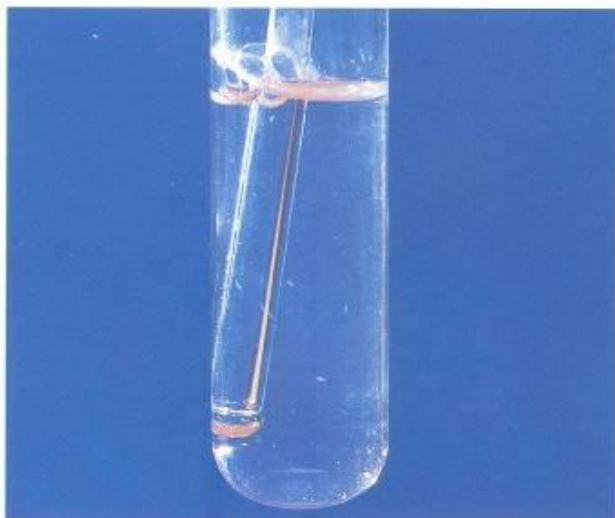
calcium carbonate + carbon dioxide + water → calcium hydrogen-carbonate



Calcium hydroxide (slaked lime) is mixed with sand to give mortar. When it is mixed with water and allowed to set a strongly bonded material is formed, which is used to hold bricks together. The hardening of mortar takes place as the following reaction occurs.

calcium hydroxide + carbon dioxide → calcium carbonate + water

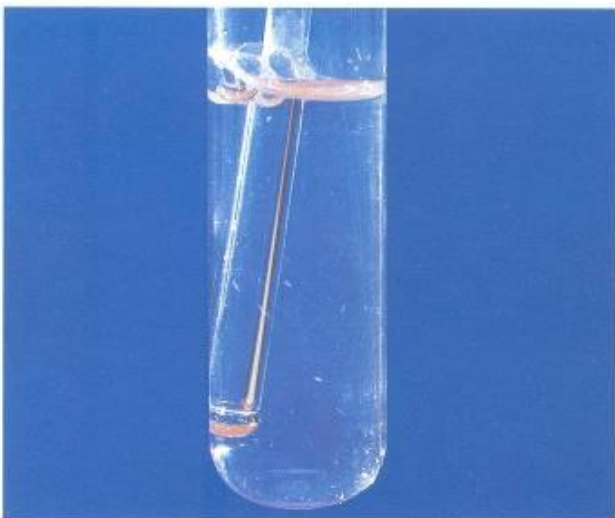




a At the start of the test.



b The limewater turns cloudy.



c If the CO_2 is bubbled for a longer length of time the white precipitate dissolves.

Figure 8.10 Testing for carbon dioxide.

Questions

1. Some suggested building alternatives to limestone are sandstone and granite. What would be the benefits of using either of these two materials instead of limestone? For further information, consult Chapter 17 as well as available Internet sites.
2. Devise an experiment that you could carry out in the laboratory to establish whether or not powdered limestone is better at curing soil acidity than calcium oxide.
3. Why is calcium oxide described as a 'base' in the reaction which occurs in question 2?

Carbonates and hydrogencarbonates

Carbonates form an important range of compounds. They are all salts of carbonic acid (H_2CO_3) and contain the carbonate ion (CO_3^{2-}). Many of them occur naturally in rock formations. For example, in addition to the calcium carbonate in limestone, chalk and marble, malachite is copper(II) carbonate and dolomite is magnesium carbonate (Figure 8.11).



Figure 8.11 Three naturally occurring carbonates – dolomite (left) with magnesite in front and malachite (right).

Properties of carbonates

Most metal carbonates thermally decompose when heated to form the metal oxide and carbon dioxide gas. For example,

copper (II) carbonate \rightarrow copper (II) oxide + carbon dioxide

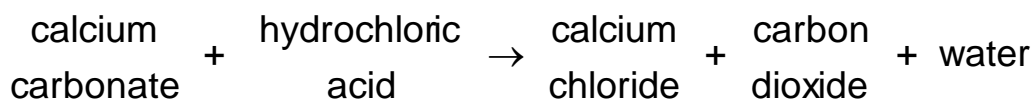


Group 1 metal carbonates, except for lithium carbonate, do not dissociate on heating. It is generally found that the carbonates of the more reactive metals are much more difficult to dissociate than, for example, copper(II) carbonate.

Carbonates are generally insoluble in water except for those of sodium, potassium and ammonium.

Carbonates react with acids to form salts, carbon dioxide and water

(Chapter 7). For example, calcium carbonate reacts with dilute hydrochloric acid to form calcium chloride, carbon dioxide and water.



This reaction is used in the laboratory preparation of carbon dioxide. It is also used as a test for a carbonate because the reaction produces carbon dioxide which causes effervescence and if bubbled through limewater turns it chalky white.

Sodium carbonate - an important industrial chemical

Approximately 1 million tonnes of sodium carbonate are manufactured in the UK annually, mainly by the Solvay process. Sodium carbonate is used for glass manufacture, brine treatment and water purification, as well as in the manufacture of detergents and textiles (Figure 8.12).

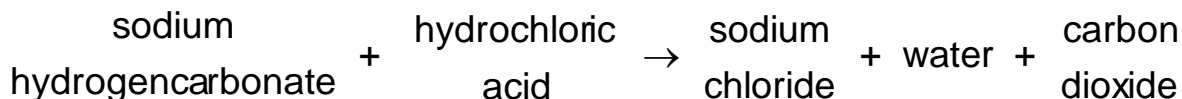


Figure 8.12 Sodium carbonate is an important chemical. It has many uses, for example in the manufacture of glass.

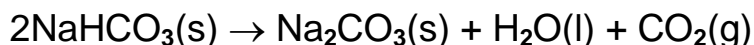
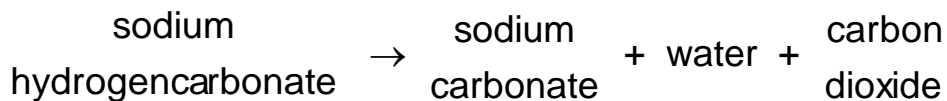
Hydrogencarbonates

Almost all the hydrogencarbonates are known only in solution because they are generally too unstable to exist as solids. They contain the hydrogencarbonate ion (HCO_3^-).

Sodium hydrogencarbonate (NaHCO_3) is the most common solid hydrogencarbonate. It is found in indigestion remedies because it reacts with the excess stomach acid (HCl) which causes indigestion.



Sodium hydrogencarbonate is the substance which is put into 'plain' flour to make it into 'self-raising' flour. When heated, the sodium hydrogencarbonate in the flour decomposes to give carbon dioxide gas, which makes the bread or cakes rise.

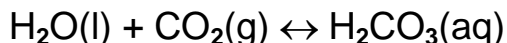


Calcium hydrogencarbonate and magnesium hydrogencarbonate are two of the hydrogencarbonates found only in solution. They are the substances mainly responsible for the hardness in water.

Hardness in water

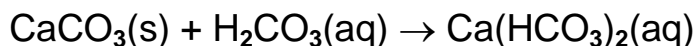
Rainwater dissolves carbon dioxide as it falls through the atmosphere. A small fraction of this dissolved carbon dioxide reacts with the water to produce carbonic acid, which is a weak acid.

water + carbon dioxide \leftrightarrow carbonic acid



As this solution passes over and through rocks containing limestone and dolomite, the weak acid in the rain attacks these rocks and very slowly dissolves them. The dissolved substances are called calcium and magnesium hydrogencarbonates.

calcium carbonate + carbonic acid \rightarrow calcium hydrogencarbonate



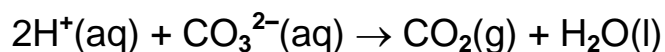
Some of the rock strata may contain gypsum (calcium sulphate, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), anhydrite (CaSO_4) or kieserite ($\text{MgSO}_4 \cdot \text{H}_2\text{O}$), which are very sparingly soluble in water. The presence of these dissolved substances causes the water to become **hard**. Hardness in water can be divided into two types — **temporary** and **permanent**. Temporary hardness is so called because it can be removed easily by boiling. Permanent hardness is so called because it is much more difficult to remove and certainly cannot be removed by boiling.



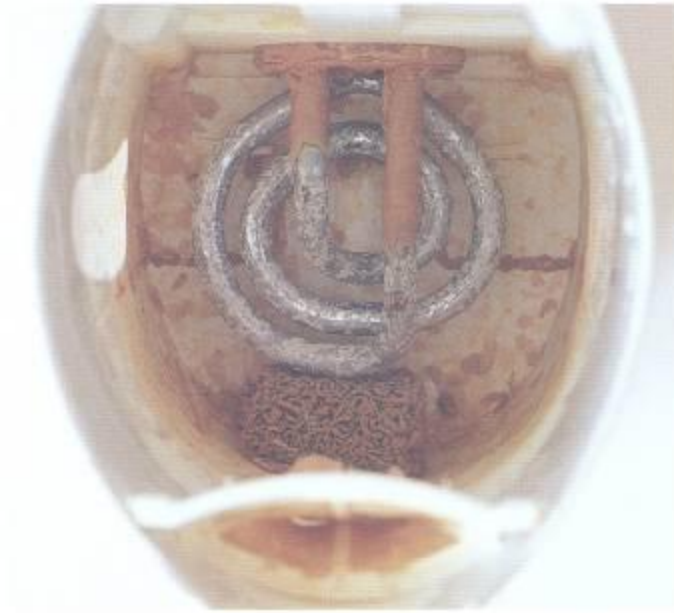
Figure 8.13 Evaporating pure water leaves no deposit, while temporary or permanent hard water leaves white deposits behind.

Temporary hardness is caused by the presence of dissolved calcium or magnesium hydrogencarbonates. Permanent hardness is caused by the presence of dissolved calcium or magnesium sulphates. When water containing these substances is evaporated, a white solid deposit is left behind (Figure 8.13). These white deposits are calcium or magnesium sulphates and/or calcium carbonate. Calcium carbonate causes the 'furring' in kettles that occurs in hard water areas (Figure 8.14a). This furring may be removed by the addition of a dilute acid:

hydrogen ions + carbonate ion \rightarrow carbon dioxide + water



Also, blockages in hot water pipes are caused by a similar process to the furring of kettles. A thick deposit of limescale builds up.



a The deposit in kettles ('furring') is caused by calcium carbonate from hard water.



b The thick deposits of limescale in the pipe have resulted from heating hot water.

Figure 8.14

Stalactites and stalagmites are found in underground caverns in limestone areas. They are formed from the slow decomposition of water containing calcium or magnesium hydrogencarbonates (Figure 8.15).

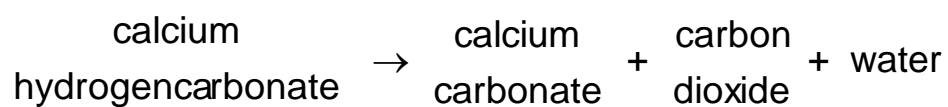




Figure 8.15 Stalactites and stalagmites have formed over hundreds of thousands of years.

Effect of hard water on soap

If you live in one of the hard water areas shown in Figure 8.16 then you may have noticed that it is difficult to make the soap lather. Instead, the water becomes cloudy. This cloudiness is caused by the presence of a solid material (a precipitate) formed by the reaction of the dissolved substances in the water with soap (sodium stearate). This white precipitate is known as scum (Figure 8.17).

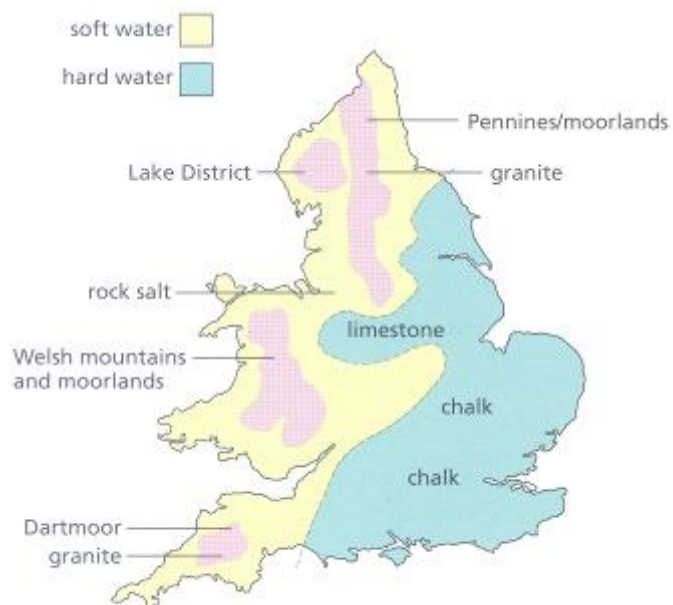
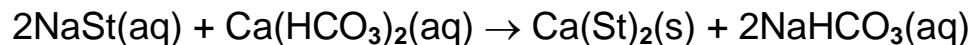
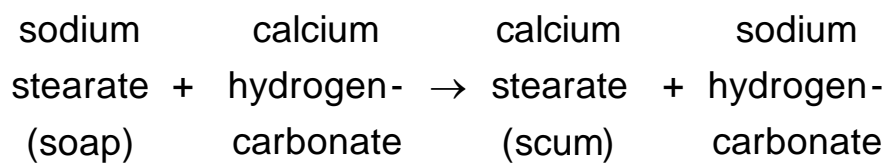


Figure 8.16 The hard and the soft water areas in England and Wales.



Figure 8.17 Soap and hard water form scum.



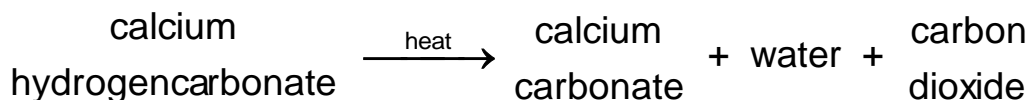


The amount of soap required to just produce a lather can be used to estimate the hardness in water.

To overcome the problem of scum formation, soapless detergents have been produced which do not produce a scum because they do not react with the substances in hard water. For further discussion of soapless detergents see Chapter 14.

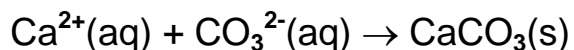
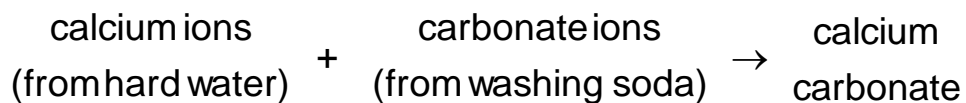
Removal of hardness

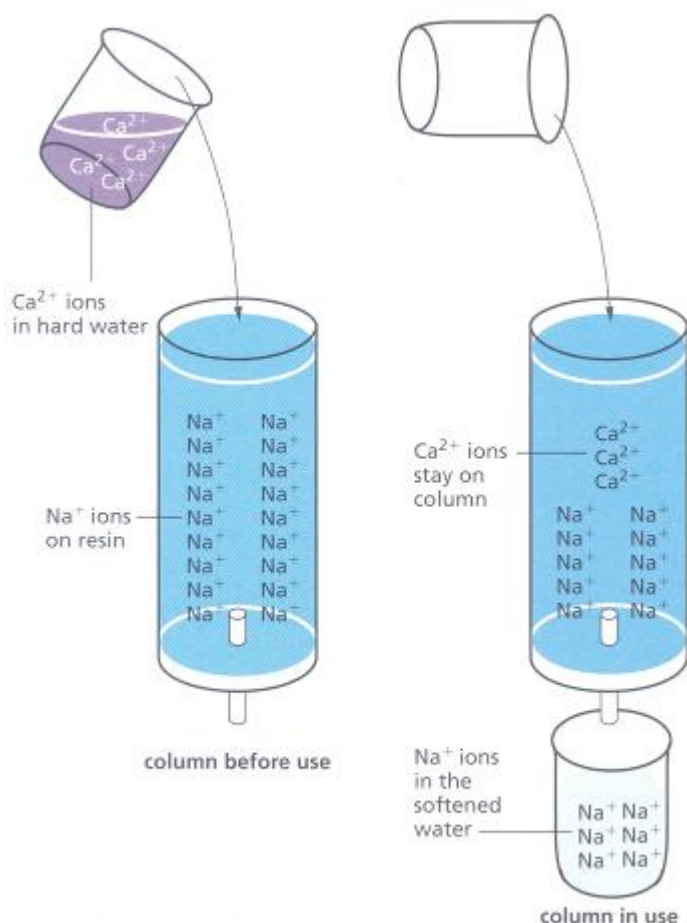
Temporary hardness is easily removed from water by boiling. When heated the calcium hydrogencarbonate decomposes, producing insoluble calcium carbonate.



The substances in permanently hard water are not decomposed when heated and therefore cannot be removed by boiling. Both types of hardness can be removed by any of the following.

- **Addition of washing soda ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) crystals.** In each case the calcium or magnesium ion, which actually causes the hardness, is removed as a precipitate and can, therefore, no longer cause hardness.





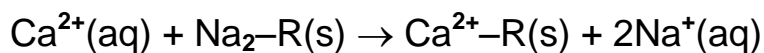
b Ion exchange resin.

Figure 8.18

a The ion exchange process.

- **Ion exchange.** The water is passed through a container filled with a suitable resin containing sodium ions. The calcium or magnesium ions causing the hardness are exchanged for the sodium ions in the resin (Figure 8.18).

calcium ion + sodium-resin \rightarrow calcium-resin + sodium ion



When all the sodium ions have been removed from the resin, it can be regenerated by pouring a solution of a suitable sodium salt through it.

- **Distillation.** The water is distilled away from the dissolved substances. This method, however, is far too expensive to be used on a large scale.

Advantages and disadvantages of hard water

As we mentioned earlier, there are some problems associated with hard water; however, there are advantages as well. Both are detailed in Table 8.1.

Table 8.1 Disadvantages and advantages of hard water.

Disadvantages	Advantages
Wastes soap	Has a nice taste
Causes kettles to fur	Calcium ions in hard water are required by the body for bones and teeth
Can cause hot water pipes to block	Coats lead pipes with a thin layer of lead (II) sulphate or lead (II) carbonate and cuts down the possibility of lead poisoning
Can spoil the finish of some fabrics	Is good for brewing beer

Questions

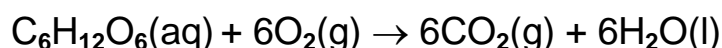
1. Write word and balanced chemical equations to show the effect of heat on cobalt carbonate (CoCO_3) and nickel carbonate (NiCO_3).
2. Make a list of the main methods of softening hard water. In each case write a suitable equation(s) to summarise the chemical reactions involved.
3. One of the substances found in some temporary hard waters is magnesium hydrogencarbonate. Write word and balanced chemical equations to show the effect of heat on this substance in aqueous solution.
4. Write word and balanced chemical equations for the reaction of dilute hydrochloric acid with:
 - a. copper (II) carbonate
 - b. sodium carbonate.

Carbon dioxide

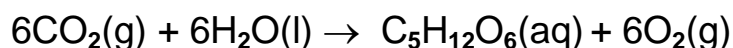
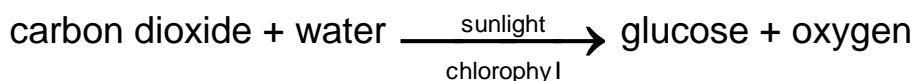
Carbon forms two oxides - carbon monoxide (CO) and carbon dioxide (CO₂). Carbon dioxide is the more important of the two, and in industry large amounts of carbon dioxide are obtained from the liquefaction of air. Air contains approximately 0.03% by volume of carbon dioxide. This value has remained almost constant for a long period of time and is maintained via the **carbon cycle** (Figure 8.19). However, scientists have recently detected a slight increase in the amount of carbon dioxide in the atmosphere.

Carbon dioxide is produced by burning fossil fuels. It is also produced by all living organisms through **aerobic respiration**. Animals and plants take in oxygen and breathe out carbon dioxide.

glucose + oxygen → carbon dioxide + water + energy



Carbon dioxide is taken in by plants through their leaves and used together with water, taken in through their roots, to synthesise sugars. This is the process of **photosynthesis**, and it takes place only in sunlight and only in green leaves, as they contain **chlorophyll** (the green pigment) which catalyses the process.



This cycle has continued in this manner for millions of years. However, scientists have detected an imbalance in the carbon cycle due to the increase in the amount of carbon dioxide produced through burning fossil fuels and the deforestation of large areas of tropical rain forest. The Earth's climate is affected by the levels of carbon dioxide (and water vapour) in the atmosphere. If the amount of carbon dioxide, in particular, builds up in the air, it is thought that the average temperature of the Earth will rise. This effect is known as the **greenhouse effect** (Figure 8.20).

Some energy from the Sun is absorbed by the Earth and its atmosphere. The remainder is reflected back into space. The energy that is absorbed helps to heat up the Earth. The Earth radiates some heat energy back into space but the 'greenhouse gases', including carbon dioxide, prevent it from escaping. This effect is similar to that observed in a greenhouse where sunlight (visible/ultraviolet radiation) enters through the glass panes but heat (infrared radiation) has difficulty escaping through the glass.

The long-term effect of the higher temperatures will be the gradual melting of ice caps and consequent flooding in low-lying areas of the Earth. There will also be changes in the weather patterns which would affect agriculture worldwide.

These problems have been recognised by nations worldwide. Recent agreements under the Kyoto Accord between nations mean that there will be some reduction in the amount of carbon dioxide (and other greenhouse gases) produced over the next few years. However, there is still a long way to go.

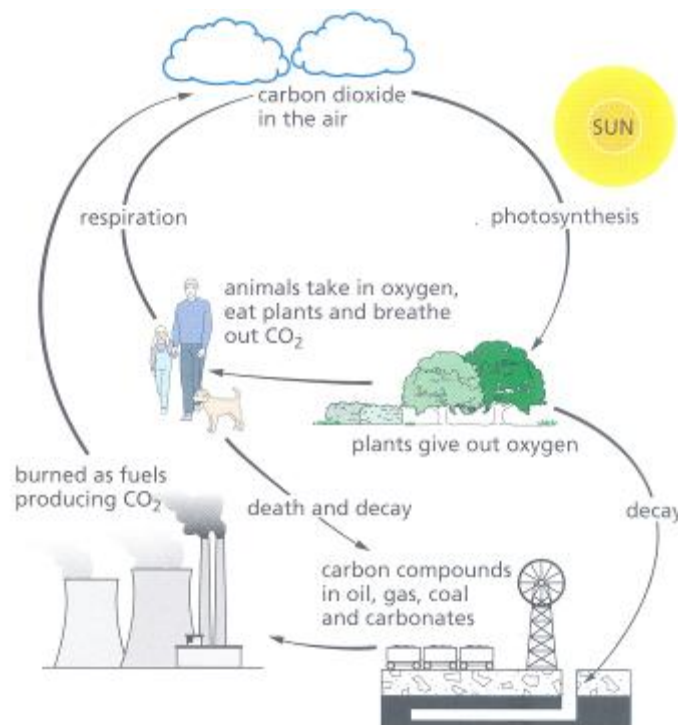


Figure 8.19 The carbon cycle.

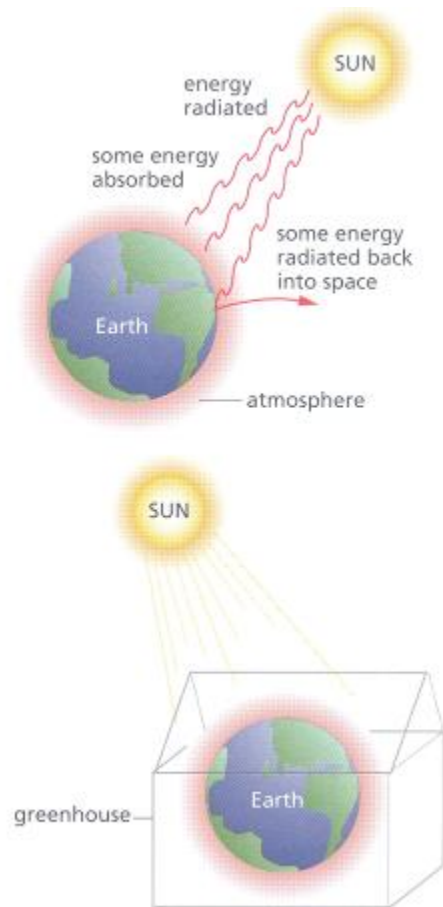
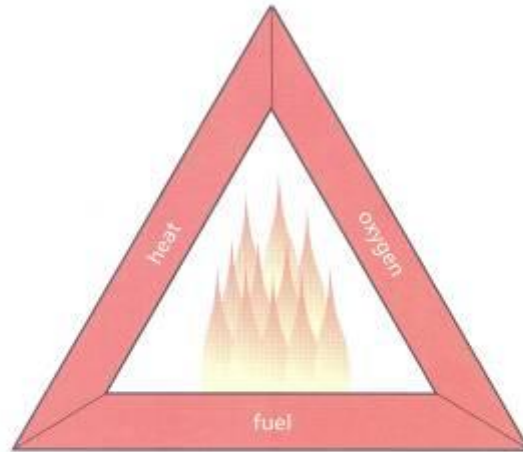


Figure 8.20 The greenhouse effect.

Uses of carbon dioxide

Carbon dioxide has some important uses.

- **Carbonated drinks.** Large quantities are used to make soda and mineral waters as well as beer. The carbon dioxide gas is bubbled into the liquid under pressure, which increases its solubility.
- **Fire extinguishers.** It is used in extinguishers for use on electrical fires. Carbon dioxide is denser than air and forms a layer around the burning material. It covers the fire and starves it of oxygen. Carbon dioxide does not burn and so the fire is put out (Figure 8.21).
- **Refrigerants.** Solid carbon dioxide (dry ice) is used for refrigerating ice cream, meat and soft fruits. It is used for this purpose because it is colder than ice and it sublimates, and so it does not pass through a potentially damaging liquid stage.
- **Special effects.** Carbon dioxide is used to create the 'smoke' effect you may see at pop concerts and on television. Dry ice is placed in boiling water and it forms thick clouds of white 'smoke'. It stays close to the floor due to the fact that carbon dioxide is denser than air.
- **Heat transfer agents.** Carbon dioxide gas is used for transferring heat in some nuclear power stations.



a Any fire needs fuel, oxygen and heat. If any of these is removed, the fire triangle is destroyed and the fire will be extinguished.

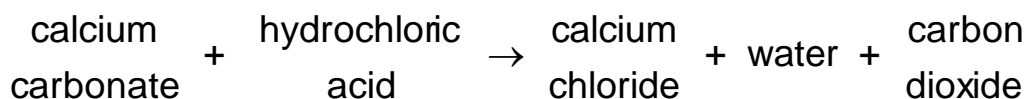


b Carbon dioxide fire extinguishers starve the fire of oxygen. They are used mainly for electrical fires.

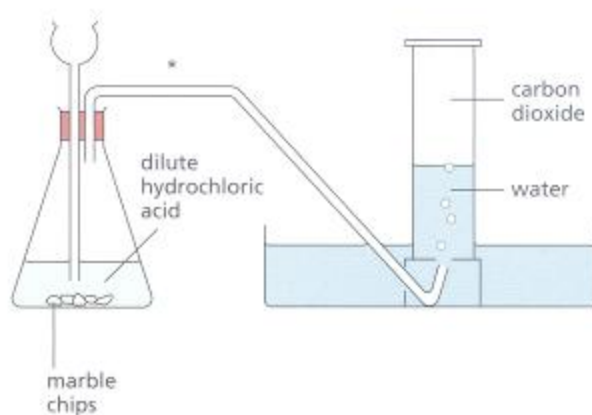
Figure 8.21

Laboratory preparation of carbon dioxide gas

In the laboratory the gas is made by pouring dilute hydrochloric acid on to marble chips (CaCO_3).



A suitable apparatus for preparing carbon dioxide is shown in Figure 8.22.



If dry gas is required, then it is passed through concentrated sulphuric acid (to dry it) and then collected as shown below

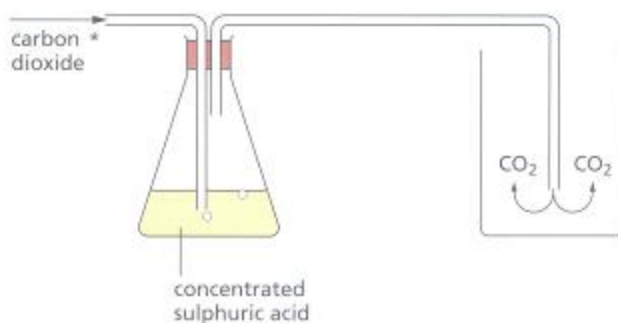


Figure 8.22 Preparation and collection of carbon dioxide gas.

Properties of carbon dioxide gas

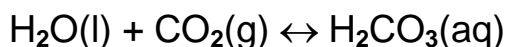
Physical properties Carbon dioxide is:

- a colourless gas
- sparingly soluble in water
- denser than air.

Chemical properties

- When bubbled into water it dissolves slightly and some of the carbon dioxide reacts, forming a solution of the weak acid carbonic acid which shows a pH of 4 or 5.

water + carbon dioxide \leftrightarrow carbonic acid



- It will support the combustion only of strongly burning substances such as magnesium. This burning reactive metal decomposes the carbon dioxide to provide oxygen for its continued burning in the gas. This reaction is accompanied by much crackling (Figure 8.23).

magnesium + carbon dioxide \rightarrow magnesium oxide + carbon

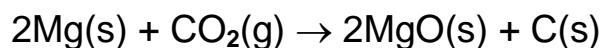
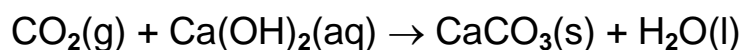
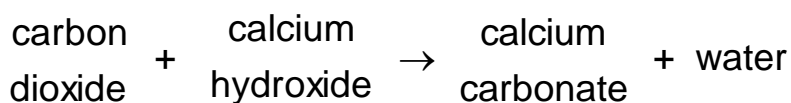




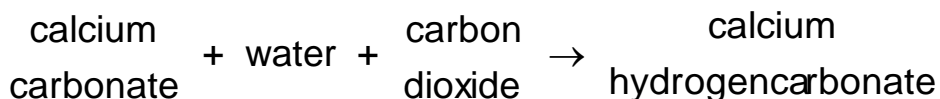
Figure 8.23 When magnesium burns in carbon dioxide gas, magnesium oxide (white) and carbon (black) are produced.

This reaction is a good example of a **redox** process (Chapter 2). Which of these substances has been oxidised and which has been reduced? Which of these substances is the oxidising agent and which is the reducing agent?

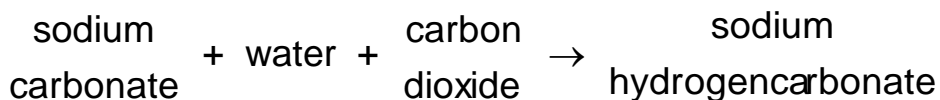
- When carbon dioxide is bubbled through limewater (calcium hydroxide solution), a white precipitate is formed. This white solid is calcium carbonate (CaCO_3). This reaction is used as a test to show that a gas is carbon dioxide.



If carbon dioxide is bubbled through this solution continuously then it will eventually become clear. This is because of the formation of soluble calcium hydrogencarbonate solution.



• Carbon dioxide reacts with strong alkalis, such as sodium hydroxide, to form carbonates. A solution of sodium hydroxide can be used to absorb carbon dioxide from the air. If excess carbon dioxide is bubbled through a solution of an alkali then a white precipitate may be obtained. When using sodium hydroxide this is due to the formation and precipitation of the sodium hydrogencarbonate:



Questions

- 1 List the important uses of carbon dioxide and for each use you have given in your answer explain why carbon dioxide is used.
- 2 Describe an experiment that you could carry out to show that carbonic acid is a weak acid.
- 3 Why is it not possible to use dilute sulphuric acid to make carbon dioxide from limestone in the laboratory?
- 4 When carbon dioxide is 'poured' from a gas jar on to a burning candle the candle goes out. What properties of carbon dioxide does this experiment show?

Checklist

After studying Chapter 8 you should know and understand the following terms.

Carbonate A salt of carbonic acid containing the carbonate ion, CO_3^{-2} , for example CuCO_3 .

Carbon dioxide A colourless, odourless gas, soluble in water, producing a weak acid called carbonic acid. It makes up 0.03% of air. It is produced by respiration in all living things and by the burning of fossil fuels. It is consumed by plants in photosynthesis.

Greenhouse effect The absorption of reflected infrared radiation from the Earth by gases in the atmosphere such as carbon dioxide (a greenhouse gas) leading to atmospheric warming.

Hardness of water This is caused by the presence of calcium (or magnesium) ions in water, which form a 'scum' with soap and prevent the formation of a lather. There are two types of hardness:

- temporary hardness — caused by the presence of dissolved calcium (or magnesium) hydrogencarbonate
- permanent hardness — this results mainly from dissolved calcium (or magnesium) sulphate.

Lime A white solid known chemically as calcium oxide (CaO). It is produced by heating limestone. It is used to counteract soil acidity and to manufacture calcium hydroxide (slaked lime). It is also used as a drying agent in industry.

Limestone A form of calcium carbonate (CaCO_3). Other forms include chalk, calcite and marble. It is used directly to neutralise soil acidity and in the manufacture of iron and steel, glass, cement, concrete, sodium carbonate and lime.

Photosynthesis The chemical process by which green plants synthesise their carbon compounds from atmospheric carbon dioxide using light as the energy source and chlorophyll as the catalyst.

Removal of hardness Temporary hardness is removed by boiling. Both temporary and permanent hardness are removed by:

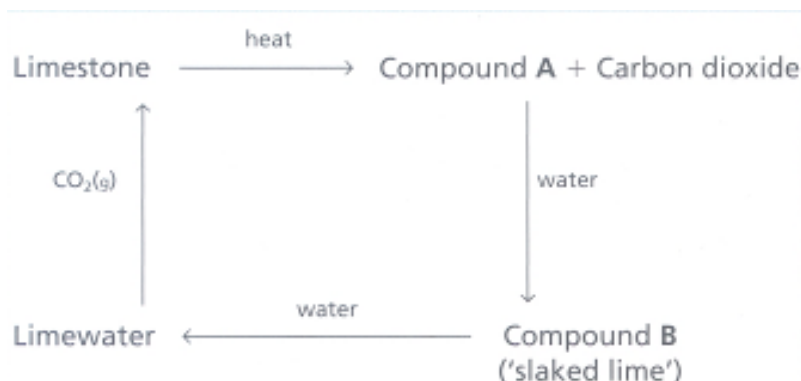
- addition of washing soda (sodium carbonate)
- ion exchange
- distillation.

Thermal decomposition The breakdown of a substance under the influence of heat.

Inorganic Carbon Chemistry

Additional questions

1. This question is about the limestone cycle.



a. Name and give the formula of:

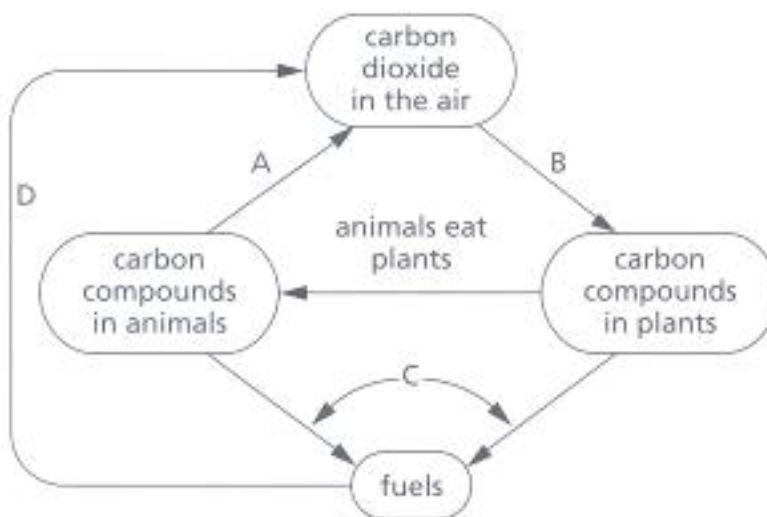
- (i) compound A
- (ii) compound B.

b. Write balanced chemical equations for the formation of both compounds A and B.

c. Name and give the symbol/formula for the ions present in limewater.

d. Describe with the aid of a balanced chemical equation what happens if carbon dioxide is bubbled through limewater until there is no further change.

2. The diagram shown below is a simplified version of the carbon cycle.



a. Name the processes A, B, C and D.

b. Write balanced chemical equations to represent the processes taking place in A and B.

c. There are a number of fuels which could be placed in the 'fuels' box. Name three such fuels.

d. 'The air contains approximately 0.03% of carbon dioxide and this amount is almost constant.' Explain why this is a true statement.

e. (i) What is the 'greenhouse effect'?

(ii) Discuss the possible consequences of the green-house effect if nothing is done to counteract it.

3. Limestone is an important raw material used in many different industries.

a. One of the properties of limestone is that it reacts with acids.

(i) Why do farmers spread powdered limestone on their fields?

(ii) How can buildings made of limestone be affected by 'acid rain'?

(iii) Write an ionic equation which would represent the reactions taking place in both (i) and (ii).

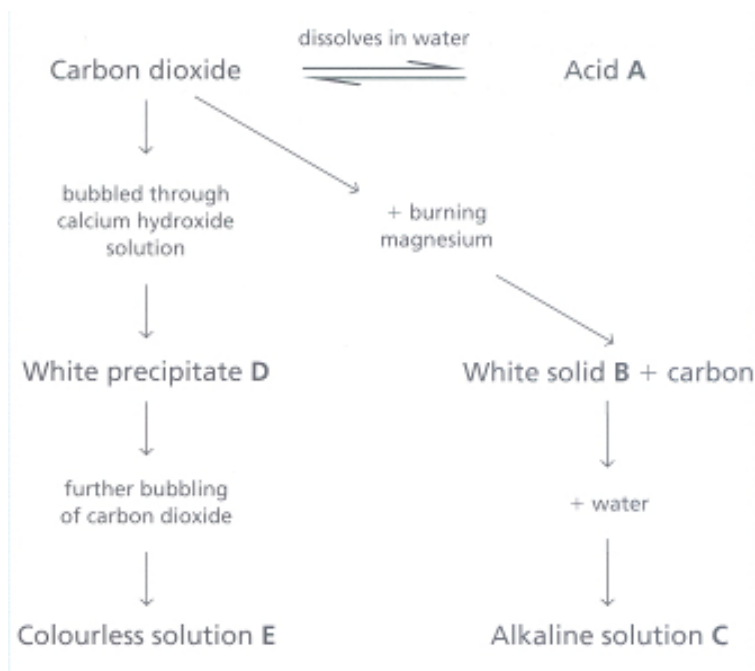
b. Limestone is used in the manufacture of iron (Chapter 9).

(i) Why is it added to the blast furnace along with coke and haematite?

(ii) Write chemical equations for the reactions it is involved in, in this process.

c. (i) Name a building material made by heating a mixture of limestone and clay in a rotary kiln.

(ii) What substance is the dry product produced in the rotary kiln added to? Explain why this substance is added.



4a. Name and give the formulae of substances A to E.

b. Write balanced chemical equations for the reactions in which compounds B, C and E were formed.

c. Where would you expect to find acid A?

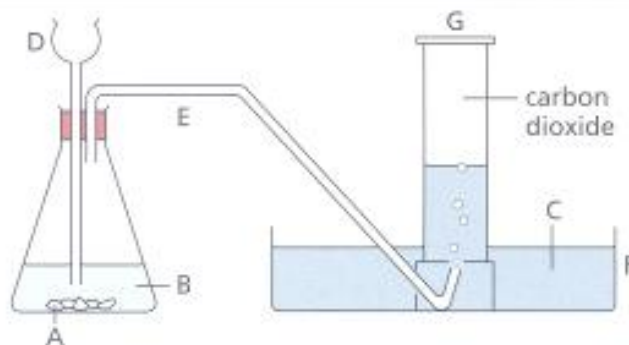
d. Universal indicator solution was added to solution C. What colour did it go?

e. Upon addition of dilute hydrochloric acid to solution C, a neutralisation reaction took place.

(i) Write a balanced chemical equation for the reaction taking place.

(ii) Name the salt produced in this reaction.

5. The following question is about carbon dioxide.



- Name and give the formula of each of the substances A, B and C.
 - Identify by name the different pieces of apparatus D, E, F and G.
 - Draw and label the apparatus that should be used if a dry sample of carbon dioxide gas was required.
 - When a gas jar containing carbon dioxide is held over a burning wooden splint and the cover removed, the flame goes out. State two properties of carbon dioxide illustrated by this observation.
 - Carbon dioxide is also produced when zinc carbonate is heated strongly.
 - Write a balanced chemical equation for the reaction taking place.
 - Name the process which is taking place as the zinc carbonate is heated.
 - Calculate the volume of carbon dioxide that would be produced (measured at room temperature and pressure (rtp)) if 12.5g of zinc carbonate were heated strongly.
- (One mole of any gas occupies 24 dm^3 at rtp. Ar: Zn = 65; C = 12; O = 16)

6. Explain the following.

- a. Industry normally requires water which has been softened.
- b. Hard water is good for the promotion of healthy bones and teeth.
- c. Hard water causes kettles to fur. This 'fur' can be removed by using a dilute acid
- d. Hard water wastes soap.
- e. Hard water can coat lead pipes and reduce the possibility of lead poisoning.

7. The results of testing five samples of water from different parts of the country are shown in the table below. The soap solution was gradually added to 25 cm³ of each sample of water with shaking until a permanent lather (one which lasts for 20 seconds) was obtained.

Water sample/25 cm ³	Volume of soap solution added/cm ³	
	Before boiling	After boiling
A	12	1
B	13	6
C	11	11
D	14	3
E	16	16

a. (i) Which samples are permanently hard?

(ii) Which samples are temporarily hard?

(iii) Which sample contains both temporary and permanent hardness?

b. Name a compound which could be present in sample D but not in sample E.

c. Name a compound which could be present in sample E but not in sample D.

d. Explain how the compound you have named in c gets into the water.

e. Sample E was distilled. The water collected was tested with soap solution. What volume of soap solution might you expect to be added to produce a permanent lather? Comment on your answer.

8. Lime (calcium oxide) is produced in very large quantities in a lime kiln. The equation for the reaction is:



a. How much limestone would be needed to produce 61.60 tonnes of lime?

(A_r : C = 12; O = 16; Ca = 40)

b. Why is the carbon dioxide swept out of the lime kiln?

c. Give three uses of lime.

d. (i) What problems are associated with the large-scale quarrying of limestone?

(ii) What steps have been taken to overcome or reduce the problems you have outlined in (i)?